IGSI

UNIT - 7

CHEMISTRY FOR CLASS IX

# CHEMICAL REACTIONS

# KINETICS

राष्ट्रीय शैक्षिक अनुसंधान और प्रशिक्षण परिषद्
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# PREFACE

The present series of twelve chemistry units has been developed for try-out of the Individually Guided System of Instruction (IGSI) in class IX The description of IGSI and first two units of this series of units are available under a separate cover. This new system of instruction and the units have been developed along the lines of the National Policy on Education (NPE-86) and involve the participation of pupils in the process of learning The units are suited for self-study with occasional help from a tutor In the present context, these units will serve as examplar self-study material for secondary stage chemistry. In developing this unit, I was assisted by some of the chemistry teachers.

This unit contains an introduction for motivation, arousing interest, and to link the present unit with preceeding and next units. The objectives given in this unit are the expected learning outcomes, so that the pupil will know the ultimate goals he has to achieve. The suggested reading material provided in the unit guides the pupil to achieve pre-stated objectives. A number of intext and post-text questions, activities, and problems have been included to provide enough practice and chance for self evaluation.

The suggestions for the improvement of this unit will be welcomed.

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May, 1988

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## III. Suggested Reading Material

7.1 Types of chemical reactions

We come across a large number of chemical changes On the basis of changes taking place in them, the chemical reactions can be classified into various categories Let us discuss these categories.

### (a) Combination reactions

The reactions in which two or more substances combine to form a compound are called combination reactions The reactants in such reactions can be either elements or compounds These can be written as

$$A + B \longrightarrow C$$

For example, many elements combine with oxygen on burning in air A few reactions of this type are given below

$$2\,Mg\,(s) + O_2\,(g) \longrightarrow 2\,MgO\,(s)$$

$$S\,(s) + O_2\,(g) \longrightarrow SO_2\,(g)$$

$$C\,(s) + O_2\,(g) \longrightarrow CO_2\,(g)$$

Some elements combine together on heating For example

$$Zn\,(s) + S\,(s) \longrightarrow ZnS\,(s)$$

$$Fe\,(s) + S\,(s) \longrightarrow FeS\,(s)$$

In some reactions two compounds combine with each other to form a new compound. For example,

$$Ca\,O\,(g) + C\,O_2\,(g) \longrightarrow CaCO_3\,(s)$$

$$NH_3\,(g) + H\,Cl\,(g) \longrightarrow NH_4\,Cl\,(s)$$

**Activity** : Combination of magnesium with oxygen.

**Procedure** : Hold a small piece of magnesium ribbon with a pair of tongs and heat it in the flame of a spirit lamp or Bunsen burner. After sometime, it starts burning. Now remove it from the flame and observe

1. Does it continue to burn ?

2 During burning, the heat is being absorbed or evolved ? Write the chemical equation for the reaction.

**Caution** : Do not look directly at the burning of magnesium. You may damage your eyes.

### (b) Decomposition reactions

These reactions are just the opposite of combination reactions. In these reactions, a single reactant breaks up into two or more similar substances which may be elements or compounds These can be written as

$$A \longrightarrow B + C$$

Many substances decompose when heated. For example, when lead (II) nitrate is heated, it breaks up into lead (II) oxide, nitrogen (IV) oxide, also called nitrogen dioxide and oxygen.

$$\underset{\text{Lead nitrate}}{2\,Pb(NO_3)_2\,(s)} \longrightarrow \underset{\text{Lead (II) oxide}}{2\,PbO\,(s)} + 4\,NO_2\,(g) + O_2\,(g)$$

Similarly copper (II) oxide (cupric oxide), lead (II) oxide and mercury (II) oxide (mercuric oxide), when heated, decompose into their constituent elements.

$$\underset{\substack{\text{Copper (II)}\\ \text{oxide}}}{2\,CuO\,(s)} \longrightarrow 2\,Cu\,(s) + O_2\,(g)$$

$$\underset{\substack{\text{Murcury (II)}\\ \text{oxide}}}{2\,HgO\,(s)} \longrightarrow 2\,Hg\,(l) + O_2\,(g)$$

### (c) Single displacement reactions

These are the reactions in which an element replaces another from its compounds The general equation for such reactions is

$$X + YZ \longrightarrow Y + XZ$$

Zinc displaces copper from copper sulphate solution

$$Zn\ (s) + Cu\,SO_4\ (aq) \longrightarrow Cu\ (s) + Zn\,SO_4\ (aq)$$

Some more examples of displacement reactions are

(a) hydrogen is displaced from $H_2SO_4$ by iron

$$Fe\ (s) + H_2SO_4\ (aq) \longrightarrow Fe\,SO_4\ (aq) + H_2\ (g)$$

(b) tin displaces hydrogen from HCl.

$$Sn\ (s) + 2\,HCl\ (aq) \longrightarrow Sn\,Cl_2\ (aq) + H_2\ (g)$$

**Activity** : Displacement of copper from copper sulphate solution.

**Procedure** : In a test tube, take 3-4 c c solution of copper sulphate and put a granule of zinc. Shake the test tube and keep it for 3-4 minutes. Observe the zinc granule carefully and explain the observation.

### (d) Double displacement reactions

These are the reactions in which two compounds exchange their parts with each other and form two different compounds. These reactions can be written as

$$AX + BY \longrightarrow AY + BX$$

(i) Precipitation of barium sulphate from a solution of barium (II) chloride is one such reactions.

$$Ba\,Cl_2\ (aq) + H_2SO_4\ (aq) \longrightarrow Ba\,SO_4\ (s) + 2\,HCl\ (aq)$$

(ii) Precipitation of copper as copper (II) sulphide.

$$Cu\,SO_4\ (aq) + H_2\,S\ (g) \longrightarrow Cu\,S\ (s) + H_2SO_4\ (aq)$$

**Activity** : Double displacement in silver nitrate and sodium chloride solutions

**Procedure** : Take two clean test tubes. Add a pinch of silver nitrate in one tube and a pinch of sodium chloride to another. Add about 2 c.c. distilled water in each test tube and shake. Now pour the solution of silver nitrate to sodium chloride solution.

Write the chemical equation for the reaction involved.

## 7.2 Oxidation and Reduction

All of us have enjoyed colourful fireworks displayed on the occasion of Deepawali. This sort of display is caused by burning of a variety of chemical compounds and metals in the powdered form. One such metal is magnesium which on burning in air produces very bright light. Now, let us understand as to what happens when magnesium is burnt in air. Magnesium on igniting combines with oxygen of air producing a white ash, magnesium (II) oxide, MgO.

$$2\,Mg + O_2 \longrightarrow \underset{\text{(white)}}{2\,MgO}$$

This process of combination of a metal with oxygen is generally termed as *oxidation* Many other elements such as iron, calcium, aluminium, copper, sodium, hydrogen, carbon and sulphur combine with oxygen in a similar way :

$$2\,Fe + O_2 \longrightarrow \underset{\text{iron (II) oxide}}{2\,FeO}$$

$$4\,Al + 3O_2 \longrightarrow \underset{\text{Aluminium (III) oxide}}{2\,Al_2O_3}$$

$$2\,Ca + O_2 \longrightarrow \underset{\text{Calcium (II) oxide}}{2\,CaO}$$

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O$$

$$C + O_2 \longrightarrow CO_2$$

$$S + O_2 \longrightarrow SO_2$$

$$2\,CO + O_2 \longrightarrow 2\,CO_2$$

So, we can define oxidation as any process where a substance combines with oxygen.

Now consider the reactions of magnesium with nitrogen, fluorine and chlorine respectively in place of oxygen.

$$3\,Mg + N_2 \longrightarrow Mg_3N_2$$
Magnesium (II) nitride

$$Mg + F_2 \longrightarrow Mg\,F_2$$
Magnesium (II) fluoride

$$Mg + Cl_2 \longrightarrow Mg\,Cl_2$$
Magnesium (II) chloride

These are also oxidation reactions but there is no oxygen in them Thus there is a need to modify the definition of oxidation.

Let us once again consider the formation of magnesium (II) oxide from a different angle Magnesium (II) oxide is an ionic compound and is represented as $Mg^{2+}O^{2-}$

$$2\,Mg + O_2 \longrightarrow 2\,Mg^{2+}O^{2-}$$

We can split this reaction into half-reactions;

$$Mg \longrightarrow Mg^{2+} + 2\,e$$

$$O + 2\,e \longrightarrow O^{2-}$$

Magnesium ion ($Mg^{2+}$) is formed when two electrons are taken away from magnesium atom; and oxide ion ($O^{2-}$) is formed when two electrons are added to the oxygen atom. On combining these two reactions we have :

$$Mg + O \longrightarrow Mg^{2+} + O^{2-}$$

In this reaction, magnesium atom transfers its two electrons to the oxygen atom and $Mg^{2+}$ and $O^{2-}$ ions are formed. (Recall the ionic bonding in Unit-5). These oppositely charged ions unite to form magnesium (II) oxide.

$$Mg^{2+} + O^{2-} \longrightarrow Mg^{2+}O^{2-}$$

Since the molecular state of oxygen is $O_2$, we do not represent it in the atomic state and show the equation as

$$2\ Mg + O_2 \longrightarrow 2\ MgO.$$

Similarly formation of ionic sodium chloride, can be represented as follows.

$$Na \longrightarrow Na^+ + e \times 2 \qquad \ldots\ldots(i)$$

$$Cl_2 + 2e \longrightarrow 2\ Cl^- \qquad \ldots\ldots(ii)$$

$$2Na + Cl_2 \longrightarrow 2\ Na^+ + 2\ Cl^-$$

The equation (i) is multiplied by 2 since we want to have the same number of electrons transferred and accepted. Now $Na^+$ ions and $Cl^-$ ions combine together and form sodium chloride, $Na^+ Cl^-$ or simply NaCl.

So, we find that oxidation reactions involve the transfer of electrons from one species to another. This is in fact important to understand the phenomenon of oxidation.

We have learnt that oxidation of magnesium involves *loss* of electrons but these electrons have been transferred to the oxygen atom. Oxygen atom has, therefore, undergone the reverse of oxidation which we shall call as *reduction process.* Clearly the two processes of oxidation and reduction are complementary and always occur together. You cannot have one without the other. The overall reaction is therefore an oxidation-reduction or simply *redox* reaction. In the modern terminology we define :

*Oxidation* as any process in which a substance loses one or more electrons.

*Reduction* as any process in which a substance gains one or more electrons.

*Oxidation-reduction or redox reaction* is used to describe any reaction involving the transfer of electrons between reactants.

*An oxidizing agent or oxidant* is an atom, molecule, or ion which accepts electrons from the other reactant.

*A reducing agent or reductant* is an atom, molecule, or ion which gives electrons to the other reactant

The terms oxidizing agent and reducing agent sometimes cause confusion because the oxidizing agent is not oxidized and the reducing agent is not reduced. So always remember that in a redox reaction (i) an oxidizing agent oxidises the other species and gets itself reduced, and (ii) a reducing agent reduces the other species and gets itself oxidized.

To sum up we may write

*Oxidizing agent—Accepts electrons—gets itself reduced*

*Reducing agent—gives away electrons – gets itself oxidized.*

Some examples of redox reactions are given below :

*Example I* : When zinc is placed in copper (II) sulphate solution, zinc displaces copper from the solution :

$$CuSO_4 + Zn \longrightarrow Cu + Zn\,SO_4$$

$$\text{or} \quad Cu^{2+}\,(aq) + Zn\,(s) \longrightarrow Cu\,(s) + Zn^{2+}\,(aq)$$

Here, zinc atoms are giving two electrons to $Zn^{2+}$ ions. In this process $Cu^{2+}$ ions are converted into Cu atoms. On the other hand, zinc atoms are converted into $Zn^{2+}$ ions The above redox reaction can be split into two half-reactions, one representing oxidation and the other reduction.

*Oxidation* : $Zn \longrightarrow Zn^{2+} + 2\,e$ (loss of two electrons)

*Reduction* : $Cu^{2+} + 2e \longrightarrow Cu$ (gain of two electrons)

Here Zn atoms are providing necessary electrons to $Cu^{2+}$ ions, and so,

*Zn is reducing agent (or reductant)*

*$Cu^{2+}$ ion is oxidizing agent (or oxidant)*

*Example 2* : When an aqueous solution of iron (III) chloride, $FeCl_3$ (yellow in colour) is acidified with HCl and tin (II) chloride, $Sn\ Cl_2$ is added, the yellow colour changes to light green due to the formation of iron (II) chloride, $FeCl_2$.

$$2\ FeCl_3 + SnCl_2 \longrightarrow 2\ FeCl_2 + SnCl_4$$

or $$Fe^{3+}\ (aq) + Sn^{2+}\ (aq) \longrightarrow Fe^{2+}\ (aq) + Sn^{4+}\ (aq)$$

On splitting we get two half-reactions

*Oxidation* : $Sn^{2+} \longrightarrow Sn^{4+} + 2e^-$ (loss of electrons)

*Reduction* : $2\ Fe^{3+} + 2e \longrightarrow 2Fe^{2+}$ (gain of two electrons)

*Oxidizing agent* : $Fe^{3+}$ ions ( $\because$ it gains electron)
(*or oxidant*)

*Reducing agent* : $Sn^{2+}$ ions ( $\because$ it loses electrons)
(*or reductant*)

The loss and gain of electrons in a redox reaction can also be represented as follows :

$$\overbrace{2\ Fe^{3+} + Sn^{2+} \longrightarrow 2\ Fe^{2+}}^{\substack{\text{Gain of two electrons} \\ \text{Reduction}}} + Sn^{4+}$$

$$Sn^{2+} \xrightarrow[\text{Oxidation}]{\text{Loss of two electron}} Sn^{4+}$$

## Questions

1 Define oxidizing and reducing agents.

2. In the redox reaction,

$$2\ Ag\ NO_3 + Zn \longrightarrow 2\ Ag + Zn\ (NO_3)_2$$

label the oxidant and the reductant.

## 7.3 Electrolysis

We have learnt that when an electric current is passed through water containing a few drops of sulphuric acid, oxygen is liberated at one electrode and hydrogen at the other Similarly, when electricity is passed through molten lead (II) bromide, $PbBr_2$, lead is formed at one electrode while a brown coloured gas, bromine, appears at another electrode

In the above two examples, positively and negatively charged ions are produced, in acidified water $H^+$ and $OH^-$ and in molten $PbBr_2$, $Pb^{2+}$ and $Br^-$ ions respectively. Presence of ions in a medium is a must when an electric current is to be passed On passing electricity, the ions move to the oppositely charged electrodes and get discharged Ionic compounds in water or without water in molten state furnish ions for the electric current to flow. This decomposition of ionic compounds by the action of electric current is called electrolysis The compounds that furnish the ions are called electrolytes The apparatus in which electrolysis is done is called a voltameter or electrolytic cell. The electrode where positively charged ions are discharged is called the cathode and is given a negative sign. The electrode where negatively charged ions are discharged is called anode and is given a positive sign.

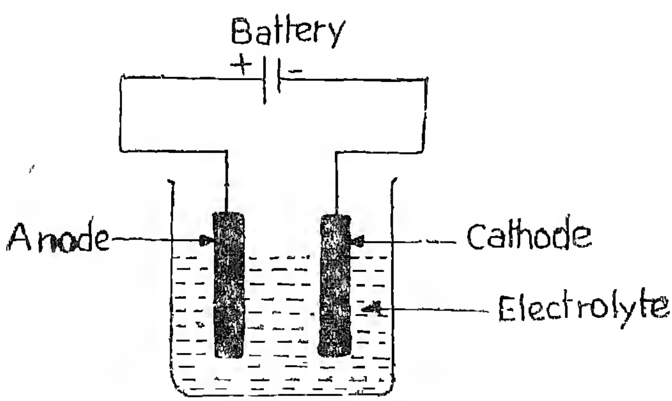


Fig 7 1 Electrolysis

## Electrolysis of lead bromide in molten state

Solid lead bromide is taken in a beaker. It is melted by heating. Heating is continued so as to keep $PbBr_2$ in molten state. On passing an electric current, lead is deposited at the cathode Brown fumes of bromine gas evolve at the anode.

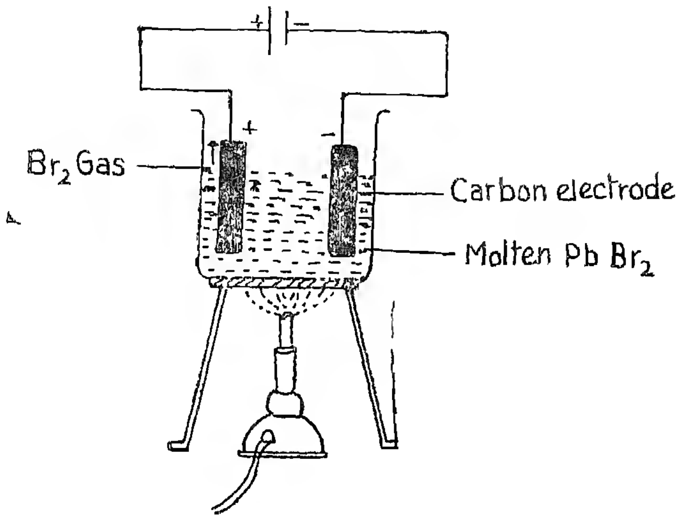


Fig. 7.2 Electrolysis of molten lead bromide

In molten lead bromide, the $Pb^{2+}$ ions and $Br^-$ ions are free to move. On passing an electric current the positive $Pb^{2+}$ ions move towards the cathode (negative) and negative $Br^-$ ions, towards the anode (positive). In an electric circuit, electrons are pushed out from the negative terminal of the cell into the cathode. From cathode these electrons are used to neutralize $Pb^{2+}$ ions.

$$Pb^{2+} + 2e \longrightarrow Pb$$

Thus, lead metal collects at the cathode. Since, $Pb^{2+}$ ions are being discharged at the cathode by gaining electrons, we call this as the reduction process.

At anode electrons are taken away from the negatively charged $Br^-$ ions. Since bromide ions lose electrons at the anode, this process is oxidation.

$$2\,Br^- \longrightarrow Br_2 + 2e \qquad \text{(at anode)}$$

## Electrolysis of aqueous solutions

We have seen that ionic compounds conduct electricity only when ions are free to move. As an alternative to melting the compound, freely moving ions may also be obtained by dissolving the compound in water. However, the electrolysis of ionic compounds has certain complications in the presence of water.

Pure water molecules split into $H^+$ and $OH^-$ as follows.

$$H_2O\,(l) \longrightarrow H^+\,(aq) + OH^-\,(aq)$$

Out of every $10^7$ (ten million) water molecules in pure sample, only one molecule splits into ions However, when ionic compounds are dissolved in water they conduct electric current very easily. To illustrate this, we will take the example of sodium chloride dissolved in water.

When sodium chloride is dissolved in water there are four different ions present in the liquid :

$Na^+$ and $Cl^-$ (from NaCl), $H^+$ and $OH^-$ (from water)

On passing an electric current in an electrolytic cell with carbon electrodes, both $Na^+$ and $H^+$ ions move towards the cathode The $OH^-$ and $Cl^-$ ions move towards anode. These ions get discharged at the respective electrodes.

You may perform any one of the following two activities.

### Activity : Electrolysis of aqueous potassium iodide solution

Set up an electrolytic cell in a beaker with carbon electrodes. Pour some 5% solution of KI Pass electric current from two fresh dry cells

(1.5 V, each) connected in series and also mark the two electrodes as cathode and anode. You will observe that brown colour will develop around the anode. This is iodine gas, $I_2$. Hydrogen gas will be released at the cathode. Write down the reactions taking place at the anode and at the cathode

With the help of a dropper take out a small amount of brown coloured solution around the anobe and mix it with a small amount of starch solution taken in a test tube Record the observations, especially the change in colour. For explanation consult your tutor/teacher.

**Activity : Electrolysis of copper sulphate solution**

Set up an electrolytic cell in a beaker with 5% solution of copper sulphate as an electrolyte and two inert carbon electrodes

(i) Pass an electric current using a 3.0 volt battery. Write the reactions occurring at cathode and anode.

(ii) Scratch the metal and collect in a test tube. Add 5 c.c. dil. $H_2So_4$ and a drop of $HNO_3$ Warm and cool What is the colour of the solution you get ? Add ammonia solution and see what happens to the original colour. (For expanation consult your teacher).

Q An aqueous solution of copper sulphate is electrolysed using carbon electrodes Name the electrode at which copper will collect. Write the reactions taking place at anode and cathode.

**Application of Electrolysis**

**1. Purifications of metals**

Electrolysis is used to purify metals. The Metal to be purified is made anode and a thin sheet of the same metal in pure state is made cathode. The electrolyte is an aqueous solution containing the cation of the

metal to be purified When an electric current is passed, the metal ions move from the anode and deposit at the cathode.

In the purification of copper, impure copper, is made anode and a thin sheet of pure copper is made cathode. Reactions occurring at cathode and anode are

$$Cu\,(s) \longrightarrow Cu^{2+}\,(aq) + 2e \qquad \textit{At Anode}$$

$$Cu^{2+}\,(aq) + 2e \longrightarrow Cu\,(s) \qquad \textit{At cathode}$$

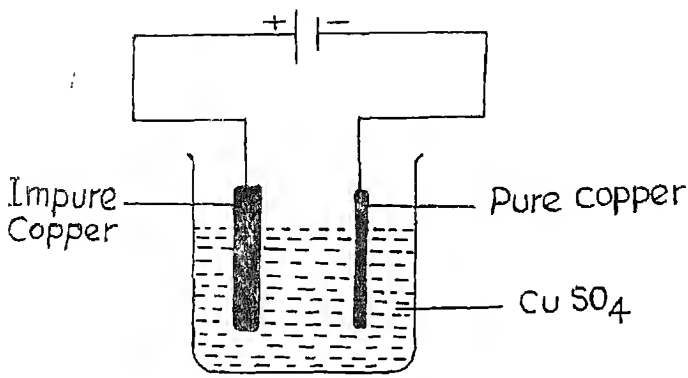


Fig. 7.3 Purification of copper

### 2. Electroplating

The technique of electrolysis, enables, thin coating of metals such as silver, gold, chromium, nickel, to be deposited on other metals such as iron and copper. In this way a protective layer of non-corrosive metal is deposited over the surface of the metal to be protected. The object to be electroplated is made the cathode.

## 7.5 Exothermic and Endothermic Reactions

In this unit, we have already discussed the burning of magnesium ribbon in air. To start this reaction, the ribbon has to be heated

initially. But once it starts burning, it continues to do so without further heating Not only this, while burning it gives out a lot of heat and light

Burning of candle is another example where chemical change is accompanied with evolution of heat.

Can you think of some more chemical processes in which heat or light or both are evolved ?

*Such reactions which are accompanied with release of energy, mostly as heat, are called exothermic reactions*

Some examples of exothermic reactions are .

(i) Neutralization reactions

$HCl\,(aq) + NaOH\,(aq) \longrightarrow NaCl\,(aq) + H_2O\,(l)$

$H_2SO_4\,(aqa) + KOH\,(aq) \longrightarrow K_2SO_4\,(aq) + 2H_2O\,(l)$

(ii) Synthesis of ammonia

$N_2(g) + 3H_2\,(g) \xrightarrow{\text{Catalyst}} 2NH_3\,(g)$

In fact, most of the reactions are exothermic In contrast, some reactions would take place only when energy is continuously supplied. For example, synthesis of nitrogen (II) oxide (nitric oxide)—

$N_2 + O_2 \longrightarrow 2NO$

*Such reactions in which energy is absorbed are called endothermic reactions.*

Some more examples of endothermic reactions are :

(i) Thermal decomposition of gypsum or calcium (II) sulphate

$CaSO_4(s) \longrightarrow CaO(s) + SO_3(s)$

(ii) Dissociation of hydrogen bromide

$2HBr\,(g) \longrightarrow H2(g) + Br_2(g)$

(iii) Dissociation of ammonia

$$2NH_3(g) \longrightarrow N_2(g) + 3H_2(g)$$

In unit 5, we have studied that various atoms in a molecule are linked through chemical bonds Whenever a bond is formed, energy is released and in the opposite change i.e breakage of a bond, energy is consumed. Whenever a chemical change takes place, some or all the bonds in the reactant molecules are broken and some new bonds are formed to make the product molecules.

In exothermic reactions, the energy consumed for bond breaking is less than the energy released during formation of new bonds. Hence in such cases, there is a net release of energy. Explain, why is energy absorbed in endothermic reactions ?

## 7.6 Rate of Reaction

Let us recall some of the changes occurring around us. We see that house hold articles made of iron get rusted gradually. This is more so during rainy season. This process is so slow that the change is visible only after a long time. Curdling of milk takes few hours. A candle burns out completely in a few hours time. These are two examples of moderate reactions. Some of the fire works (phooljhari and Anaar) give us pleasure for a few minutes.

Whereas other fire works explode in a moment only. The chemical reactions occurring here are fast. It would be interesting to consider the reaction rates in quantitative terms rather than using qualitative terms like fast, moderate and slow ?

As a reaction proceeds, the reactants are consumed and products are formed with passage of time. In other words, the concentration of reactants decreases and those of products increases with time. These changes in concentrations of reactants and products can be measured after regular intervals of time. From such data we can calculate the changes in concentrations per unit time. This change in the concentration of any of the reactant or product per unit time is known as rate of

the reaction. Let us consider a reaction in which one mole of each of the reactants react to form one mole of each of the products. For example, dissolution of iron filings in dilute sulphuric acid;

$$Fe\ (s) + H_2SO_4\ (aq) \longrightarrow Fe\ SO_4\ (aq) + H_2\ (g)$$

In one experiment, 0.1 gram iron filings were placed in 1 0 c.c. of 1.0 molar sulphuric acid. Concentrations of sulphuric acid and iron (II) sulphate in the solution were measured and recorded at regular intervals of time. (Table 7.1).

Table 7 1

| Time (Min ) | Concentration of $H_2SO_4$ (Mol/litre) | Concentration of $FeSO_4$ (Mol/litre) |
|---|---|---|
| 0 | 1.00 | 0 00 |
| 10 | 0 70 | 0 30 |
| 20 | 0.49 | 0.51 |
| 30 | 0.34 | 0.66 |
| 40 | 0.24 | 0.77 |

In the first 10 minutes, the change in concentration of sulphuric acid (a reactant)

$$\triangle CH_2\ SO_4 = (0.70-1.00)\ \text{mol litre}^{-1} = -0\ 30\ \text{mol litre}^{-1}$$

So, we see that the concentration of sulphuric acid decreases with time. In fact concentration of any reactant in any reaction decreases with time.

In other words, change in concentration of $H_2SO_4$ is negative. Mathemetically the rate of the reaction is given by the following equation;

$$\text{Average Rate} = -\frac{\Delta C_R}{\Delta t}$$

Where $\Delta C_R$ is the change (decrease) in concentration of reactant during time interval $\Delta t$.

In the example of $H_2SO_4$, after ten minutes of reaction.

$$\Delta t = (10 \text{ min} - 0 \text{ min}) = 10 \text{ min.}$$

$$\text{Hence Average Rate} = -\frac{(-0\ 3) \text{ mol litre}^{-1}}{10 \text{ min.}} = 0.03 \text{ mol litre}^{-1} \text{ min}^{-1}$$

Now let us calculate the average rate using the concentration of iron (II) sulphate.

$$Cp = \Delta C_{FeSO_4} = (0\ 30\text{-}0.00) \text{ mol litre}^{-1} = 0.30 \text{ mol liter}^{-1}$$

$$\text{Average Rate} = \frac{\Delta C_P}{\Delta t}$$

$$= \frac{0.30 \text{ mol litre}^{-1}}{10 \text{ min.}} = 0\ 03 \text{ mol litre}^{-1} \text{ min}^{-1}$$

Calculate the average rate of this reaction in time intervals (i) 10 to 20 min. (ii) 20 to 30 min and (iii) 30—40 min. using the concentration of (i) sulphuric acid and iron (II) sulphate.

You will find that, the average rate of the reaction during any interval of time is the same whether you use the changes in concentration of the reactants or products.

Now Consider another reaction

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

In this, 1 mole of $N_2$ reacts with 3 moles of $H_2$ with the formation of 2 moles of $NH_3$

Mathematically the rate of this reaction with respect to all the three species is denoted as follows :

$$-\frac{\Delta[N_2]}{\Delta t}, \quad -\frac{\Delta[H_2]}{\Delta t} \text{ and } \frac{\Delta[NH_3]}{\Delta t}$$

Where $\Delta[N_2]$ and $\Delta[H_2]$ are the decrease in concentration of $N_2$ and $H_2$ respectively in time $\Delta t$. and $\Delta[NH_3]$ is the increase in concentration of $NH_3$ in time $\Delta t$.

From the chemical equation we can see that for three moles of $H_2$ and one mole of $N_2$ are consumed and 2 moles of $NH_3$ are formed. Thus the three terms used above for denoting the rate are not equal. These can be made equal by dividing them with the respective numerals appearing in the chemical equation

Therefore

$$-\frac{\Delta[N_2]}{\Delta t} = -\frac{1}{3}\frac{[\Delta H_2]}{\Delta t} = \frac{1}{2}\frac{\Delta[NH_3]}{\Delta t}$$

## 7.7 Factors affecting Rates of Reactions

In the last section, we have studied that different reactions have different rates. We also know that food can be preserved for a long time by keeping it in refrigerator. Certain chemical reactions take place during the food spoilage. The rates of such reactions are slowed down by lowering the temperature of food in refrigerator. It shows that rate of a reaction depends on temperature. Let us study various factors on which the rate of a reaction depends.

### (a) Concentration

Let us again consider the data given in Table-7 for the reaction between iron filings and sulphuric acid. The data shows that as reaction proceeds, the concentrations of reactants decrease and those of products increase. Let us perform the following experiment.

*Experiment*. Take three boiling tubes and number them as 1,2 and 3. Take 0 5 c c conc HCl in each of them Add 4c.c water in test tube No. 1, 8 c c in no 2 and 16 c c in no 3 In each tube, add about 0 2g zinc powder The evolution of hydrogen will begin in each test tube

*Observe* the rate of evolution of hydrogen in each tube and relate it to concentration of HCl in them You will find that the rate of evolution of hydrogen is maximum in test tube-1 and minium in tube No. 3.

This experiment shows that the rate of a reaction decreases with decrease in concentration of reactants. Let us try to explain this dependance of rate of a reaction on concentration

Consider a reaction $A + B \longrightarrow C + D$.

The reaction takes place only when the rate of a reaction, depends upon the number of collisions between A and B molecules occurring per unit time. The number of such collisions depends upon the number of such molecules present This can be understood from the following diagrams

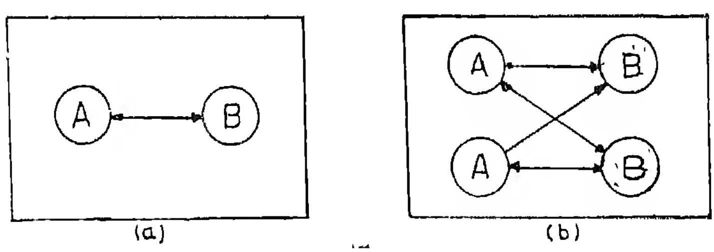


Fig. 7.4 Molecular collisions depend on number of molecules

Box (a) contains one molecule each of compounds A and B In this there is only one chance of collision Box (b) contains two molecules each of A and B i e the concentrations of A and B are twice as in the box (a). Here each molecule of A has two chances of colliding with two molecules of B In all, there are four chances of collisions between

molecules of A and B Thus in box (b) where concentrations of both the reactants A & B are more, the number of collisions resulting in chemical reaction also increases Thus the rate of reaction increases with concentration of reactants.

### (b) Temperature

Rates of most chemical reactions increase with rise in temperature This can be easily seen by heating a mixture containing zinc powder and dilute hydrochloric acid used in the previous experiment You will observe that on heating the rate of evolution of hydrogen increases. We make use of this fact on many occasions in our daily life to increase or decrease the rate of chemical reactions. For example, in winters the curdling of milk is not easy Usually before mixing curd, the milk is heated and then the container is either wrapped in a cloth and kept in tin containing wheat flour to avoid quick cooling. Heating of milk increases the rate of chemical reactions taking place during curdling.

### (c) Catalyst

**Activity** : In a test tube, take 2c c water, dissolve 0 1g oxalic acid in it. Acidify this solution with 2c.c dilute sulphuric acid & add 1 drop $KMnO_4$ solution. You will observe that at room temperature, rate of this reaction is extremely slow. Pink colour of the solution due to $KMnO_4$ is not discharged even on keeping the tube for a few minutes Now add a crystal of manganese sulphate and shake the test tube Immediately the colour of solution gets discharged Add few more drops of potassium permanganate solution to the test tube. You will observe that the colour appears immediately. Why is it so ?

In this experiment, we observe that addition of small quantity of manganese sulphate increases the rate of the reaction. Such substances which alter the rate of the reaction without undergoing a chemical change themselves are called catalysts Most of the catalysts increase the reaction rates.

Catalysts are used extensively in the chemical industry Their use, often, makes it possible to carry out the reactions at comparatively low temperatures and this makes the process cheaper.

### Examples of catalysts

1. Synthesis of ammonia by Haber's process:- Finely divided iron and molybednum are used as calatysts

$$N_2 (g) + 3 H_2 (g) \xrightarrow{Fe/Mo} 2 NH_3$$

2. Manufacture of sulphuric acid by contact process :- Platinized asbestos or vanadium pentoxide is used as catalyst.

$$2 SO_2 + O_2 \xrightarrow[V_2 O_5]{pt/Asbestos\ or} 2 SO_3$$

Catalysts are of extreme importance for the proper functioning of the human body and other biological systems. In human body, catalysts called enzymes, cause many reactions to take place rapidly at body temperature. The same reactions proceed very slowly in absence of catalysts.

### Physical state of reactants

**Activity :** Take three test tubes and label them as A,B & C. Take 2c.c of 0.5M $H_2SO_4$ in each of them. Put a strip of zinc in test tube A (If it is not available, zinc strip can be obtained from a used dry cell), zinc granules in test tube B, and zinc powder in test tube C. Now observe the rate of evolution of hydrogen in all the three test tubes.

In this experiment you will observe that reaction rate is maximum in the test tube containing zinc powder, followed by granulated zinc. The reaction is slowest in test tube containing zinc strip. From this experiment we can conclude that in case of solid reactants, the chemical activity is enhanced in the finer state of division.

The reaction rate also depends on the chemical nature of reactants.

## IV Home Assignment

1 Name the types of reactions :

(a) $AgNO_3 + NaCl \longrightarrow AgCl + NaNO_3$

(b) $PCl_5 \longrightarrow PCl_3 + Cl_2$

(c) $Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$

(d) $NH_4NO_3 \xrightarrow{\text{heat}} N_2O + 2H_2O$

(e) $H_2 + Cl_2 \longrightarrow 2HCl$

2. Show that oxidation and reduction occur simultaneously in a redox reaction. Give one example.

3. Name the substances (i) Oxidised, (ii) reduced, (iii) oxidising agent and (iv) reducing agent in the following reactions :

(i) $2Al + 3Cl_2 \longrightarrow 2AlCl_3$

(ii) $HgCl_2 + Hg \longrightarrow Hg_2Cl_2$

(iii) $2FeCl_3 \longrightarrow 2FeCl_2 + Cl_2$

4 List the endothermic and exothermic processes which take place in your surrounding or at your home

5. Will the following operations increase or decrease the rate of a reaction ?

(i) increase the temperature of the reactants

(ii) dissolving the reactants in a common solvent

(iii) increase in the concentration of reactants and

(iv) use of a catalyst.

6. Explain the following :

(a) electrolysis is a type of redox reaction.

(b) cathodic reduction is utilised for manufacture of metals.

(c) non-metals are manufactured by anodic oxidation.

## V Self Assessment

### 1 Fill in the Blanks

(1) The reaction between hydrogen and chlorine to produce hydrogen chloride gas is————————reaction.

(2) $Zn(S) + H_2SO_4\ (aq) \rightarrow ZnSO_4\ (aq) + H_2$ is a————————reaction

(3) In an exothermic reaction, heat is————————————.

(4) In an endothermic reaction, heat is———————————.

(5) Sodium chloride is——————————————electrolyte

(6) Sugar solution——————————————electric current.

(7) The rate of reaction is defined as————————in the concentration of reactants per unit time.

**2.** Identity the correct answer by putting tick mark in the following.

(1) $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ represents

(a) combination reaction (b) oxidation (c) displacement

(2) $AgNO_3\ (aq) + NaCl\ (aq) \rightarrow AgCl\ (s) + NaNO_3\ (aq)$ is

(a) redox reaction (b) double decomposition (c) exothermic

(3) In an exothermic reaction

(a) heat is absorbed (b) heat is evolved (c) temperature decreases.

(4) A catalyst is a substance that

(a) alters the reaction rate (b) increases the yield (c) takes part in a chemical reaction.

### 3. Select true and false statements

(1) $SO_2(g) + 1/2\ O_2(g) \rightarrow SO_3$ is an example of combination reaction

(2) In oxidation the electrons are lost by a substance.

(3) $N_2 + 3H_2 \longrightarrow 2NH_3(g) + \text{Heat}$, it is an endothermic reaction

(4) Copper is reduced in the reaction
$CuO(s) + H_2(g)\ Cu(s) + H_2O(l)$

(5) $CaCO_3(s) \xrightarrow{\text{heat}} CaO(s) + CO_2(g)$ is a thermal decomposition

(6) Weak electrolytes are completely dissociated in aqueous solution.

(7) Electrolytes Conduct electricity in solution and in molten state

(8) In the electrolysis of water, $H_2$ gas is liberated at cathode.

**4. Answer the following questions (In one word)**

(1) In the reaction $Cu^{2+} (aq) + Zn (s) \longrightarrow Cu (s) + Zn^{2+} (aq)$ which one is reduced

(2) What is the name of reaction in which two substances exchange ions

(3) With the increase of temperature, the rate of reaction will increase or dicrease.

(3) Which will be faster reaction

chalk + dil HCl Or chalk + conc HCl.

(5) During electrolysis which type of reaction takes place at anode?

**5. Answer the following questions**

(1) Classify the following as strong and weak electrolyte

NaCl, $CaCl_2$, $H_2CO_3$, $H_2SO_4$, $CH_3$ COOH, $HNO_3$, $CuSO_4$

(2) What is a redox reaction ?

(3) Define Rate of reaction.

# Teacher's Guide

This unit deals with the types of chemical reactions and kinetics of chemical reactions. The classification of reactions into four different types has been done on the basis of the mode in which the reactants react with one another to form products. The students may be allowed to carry out at least one experiment of each type of reactions However, if the facilities in the laboratory are not adequate, the teacher may show them experiments as class-room demonstrations The reactions may be discussed in the class with the help of experiments suggested in the unit. During the discussion, it may be emphasized that redox reactions are not a separate category of chemical reactions. The reactions like combination, decomposition and displacement may also involve oxidation and reduction processes. Suitable examples given in the unit may be taken to explain this point.

*Electrolysis* :—The phenomenon of electrolysis may be demonstrated in the class by using the solution of copper chloride as electrolyte. The carbon rods of the used battery cells are used as electrodes. The dry cells are used as the source of energy. On passing an electric current, copper gets deposited at the cathode. Chlorine gas evolves at anode and following reactions take place in the cell.

$2Cl^- \longrightarrow 2\dot{C}l + 2e$ (at anode)

$2\dot{C}l \longrightarrow Cl_2$

$Cu^{2+}(aq) + 2e \longrightarrow Cu(s)$ (at cathode)

As the two cell reactions involve electron transfers, these are redox reactions. The reaction which takes place at anode is oxidation and the one at cathode is reduction.

In discussing the rate of reaction, some examples of fast and slow reactions may be cited from daily life situations besides the examples given in the introduction of this unit. In discussing the rate of reaction, the need for dividing concentration numerals by the respective stoichiometric co-efficients appearing in the balanced chemical equation may be emphasised with the help of some more examples.

The rate of reaction depends on (i) concentration of reactants, (ii) temperature, (iii) catalyst and (iv) nature of reactants Suitable student's activities and experiments have been given in the unit to show the effects of these factors on reaction rate Depending on the facilities, the teacher may demonstrate some of the experiments in the class. The activities are meant to be performed by the students.